Topic 7: Intermolecular Forces

(a.k.a) non-bonded interactions

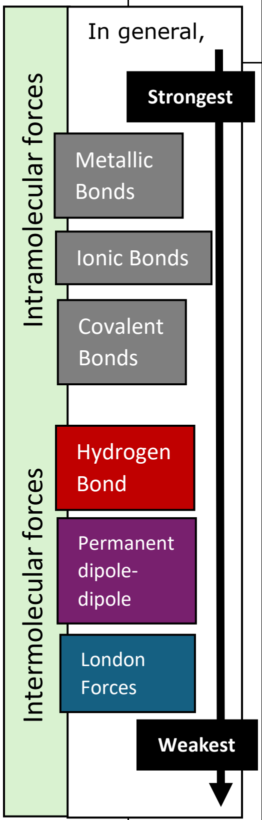
Students will be assessed on their ability to:

7.1 understand the nature of the following intermolecular forces:

i London forces (instantaneous dipole-induced dipole)

ii permanent dipole-permanent dipole interactions

iii hydrogen bonds



London (Dispersion) Forces

This (Instantaneous dipole-Induced dipole interaction) occurs because in any molecule, the electrons are constantly and randomly moving around and so its electron density fluctuates over time which results in a temporary dipole being formed. These temporary dipoles induce more dipoles in neighbouring molecules. When this happens, the δ^+ end in one molecule and the δ^- end in the neighbouring molecule are attracted to each other

They occur between all types of molecules whether polar or non-polar and between separate atoms in noble gases. They do not occur in *ionic substances*

Main Factors that affect size of London Forces

- ***Number of electrons***: The more electrons there are in a molecule, the greater the fluctuations in electron density and so the larger the instantaneous and induced dipoles created. This makes London forces stronger between the molecules and more energy will be needed to break them apart so boiling point will increase
- ***Shape & Size of molecule***: The more points of contact there are between the molecules, the greater the overall London force as instantaneous dipole-induced dipole forces exist at each point of contact between the molecules.
  That is why Long straight chain alkanes have stronger London forces as they have a larger surface area of contact between molecules compared to spherical shaped branched alkanes

Permanent dipole-dipole forces

This (Permanent dipole-Permanent dipole interaction) is stronger than London forces and so their compounds have higher boiling points.

They occur between polar molecules.

The $\delta^+$and δ^- regions of neighbouring polar molecules attract each other and hold the molecules together in a lattice-like structure.

Polar bonds form due to a difference in electronegativity

It is better to assume that the strength of each type of intermolecular force depends on whether their interaction is favorable in each case

Hydrogen Bonding

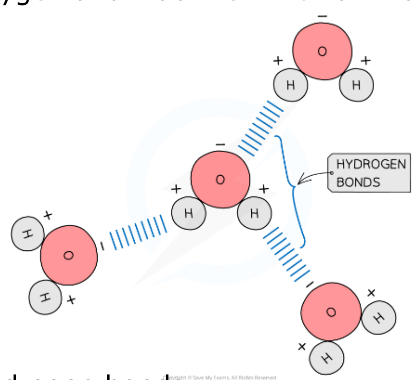
This is an intermolecular interaction (in which there is some evidence of bond formation) between a hydrogen atom of a molecule (or molecular fragment) bonded to an atom which is more electronegative than hydrogen and another atom in the same or a different molecule.

They act most often between compounds that have a hydrogen atom attached to one of the three most electronegative atoms, Nitrogen, Oxygen and Fluorine which all must have a lone pair of electrons

Note that Hydrogen bonding occurs in addition to London forces

Hydrogen Bonding through Nitrogen

- All compounds containing an -N-H group can form intermolecular hydrogen bonds
- Primary Amines can form intermolecular hydrogen bonds

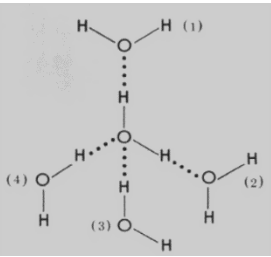
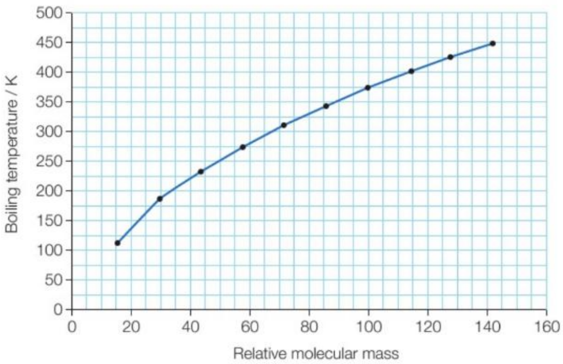


Hydrogen Bonding through Oxygen

- All compounds containing an -O-H group form intermolecular hydrogen bonds
- Alcohols form intermolecular hydrogen bonds

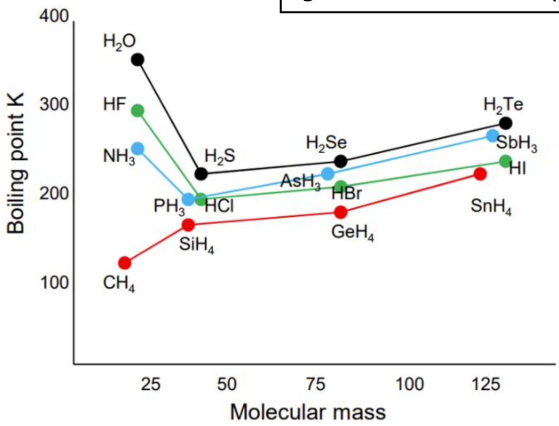
Hydrogen Bonding through Fluorine

- Only hydrogen fluoride forms intermolecular hydrogen bonds

7.2	understand the interactions in molecules, such as **H_2O, liquid NH_3 and liquid HF**, which give rise to hydrogen bonding
	Alcohols (O-H bond), ammonia (N-H bond), amines (N-H bond), carboxylic acids (O-H bond), hydrogen fluoride (H-F bond), proteins (N-H bond) can form hydrogen bonds
7.3	understand the following anomalous properties of water resulting from hydrogen bonding: i its high melting and boiling temperature when compared with similar molecules ii the density of ice compared to that of water
	1. Water can form four hydrogen bonds per molecule, because the electronegative oxygen atom has two lone pairs of electrons on it and two hydrogen atoms Thus, it can form stronger hydrogen bonds and needs more energy to break the bonds, leading to a higher melting and boiling point 2. Density of ice at 0°C is less than that of water at 0°C because the molecules in ice are arranged in rings of six, held together by hydrogen bonds. The structure creates large areas of open space and when ice melts, the ring structure is destroyed and the average distance between molecules decreases, resulting in an increase in density 
7.4	be able to predict the presence of hydrogen bonding in molecules analogous to those mentioned in 7.2
7.5	understand, in terms of intermolecular forces, physical properties shown by substances, including: i the trends in boiling temperatures of alkanes with increasing chain length ii the effect of branching in the carbon chain on the boiling temperatures of alkanes iii the relatively low volatility (higher boiling temperatures) of alcohols compared to alkanes with a similar number of electrons iv the trends in boiling temperatures of the hydrogen halides HF to HI
i	As chain length increases, 1. molecular mass increases and so does the number of electrons per molecule which causes an increase in instantaneous dipole-induced dipoles and 2. the number of contact points between molecules increases which means greater overall London forces, resulting in an increase in boiling point !Note there are only London forces in alkanes 
ii	The more branching in the molecule, the fewer the points of contact between adjacent molecules This leads to a decrease in overall intermolecular force of attraction between molecules and hence, a decrease in boiling temperature
iii	As alcohols contain an -OH group, they can form hydrogen bonds in addition to London forces. This means, compared to equivalent alkanes with similar chain length and number of electrons, the boiling point of alcohols would be higher due to the hydrogen bonding. This additional force of attraction increases the energy required to separate the molecules

ΔH_{vap} is a measure of the energy change that is required to completely separate the molecules of a liquid and convert it into a gas at the same temperature

<table>
<tr><td></td><td>A direct measure of the strength of intermolecular interactions would be the enthalpy change of vaporization
The greater the enthalpy change of vaporization, the greater the force of attraction between molecules</td></tr>
<tr><td>iv</td><td>The anomalously high boiling points of H_2O, NH_3 and HF are due to the hydrogen bonding between the molecules in addition to the London forces already present, making it require more energy to break these forces of attraction
The general increase in boiling point from HCl to HI is caused by the increasing number of electrons per molecule, thereby resulting in an increase in London forces



All hydrogen halides have London forces and permanent dipole-permanent dipole forces between its molecules
Only HF has hydrogen bonding in addition to these forces</td></tr>
<tr><td>7.6</td><td>understand factors that influence the choice of solvents, including:
i water, to dissolve some ionic compounds, in terms of the hydration of the ions
ii water, to dissolve simple alcohols, in terms of hydrogen bonding
iii water, as a poor solvent for compounds (to include polar molecules such as halogenoalkane), in terms of inability to form hydrogen bonds
iv non-aqueous solvents, for compounds that have similar intermolecular forces to those in the solvent</td></tr>
<tr><td>i</td><td>For a substance to dissolve,
• The solute particles must be separated from each other and then become surrounded by solvent particles
• The force of attraction between the solute and solvent particles must be strong enough to overcome the solvent-solvent forces and the solute-solute forces
When an ionic lattice dissolves in water, the bonds in the lattice are broken and new bonds are formed between the ions and the water molecules
The energy required to separate the ions in the solid is either completely or partially supplied by the hydration of ions
The negative ions are attracted to the δ^+ hydrogen ends of the polar water molecules
The positive ions are attracted to the δ^- oxygen ends of the polar water molecules
In general, the greater the ionic charge, the less soluble an ionic compound is</td></tr>
<tr><td>ii</td><td>The solubility of alcohols in water decreases with increasing hydrocarbon chain length as London forces predominate between the alcohol molecules</td></tr>
<tr><td>iii</td><td>Compounds that cannot form hydrogen bonds with water include
• Non-polar molecules such as the alkanes
• Some polar molecules such as ethoxyethane and halogenalkanes</td></tr>
<tr><td>iv</td><td>“Like dissolves like” meaning compounds that have similar intermolecular forces to those in the solvent will generally dissolve
Non-polar solutes will dissolve in non-polar solvents (like hydrocarbons)
Polar solutes generally dissolve in polar solvents</td></tr>
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	Further suggested practicals: i the solubility of simple molecules in different solvents ii measuring the enthalpy change of vaporisation of water iii measuring temperature changes when substances dissolve